

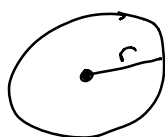
PERIODIC TRENDS:

March 26, 2018

① Atomic radius

a. Calculating the atomic radius

- Simple:



← where we find the valence e^- .

~~DONE!~~

Not quite

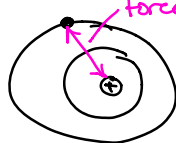
- We need to think in terms of orbitals, or electron clouds.
- "Outer-edge" of the electron cloud is not well defined.
- One solution:

Define the atomic radius ^(r) as half the distance between two adjacent nuclei of the same elem.



b. Effective nucleus charge (Z_{eff})

- Z_{eff} : The net force of attraction experienced by the valence electrons with respect to the nucleus.



- $Z_{eff} = \# \text{protons} - \text{inner electrons}$.

(activity)

c. Trends (main group elements)

- Group: - Governed by " n " the energy level
 - Adding a level of inner electrons, \therefore increases shielding of valence electrons, $\downarrow Z_{eff}$

So: Atomic radius increases as we go down a group.

- Period: - Same value of n ...
- Governed by Z_{eff} .
- As we go across a period, we keep adding 1 proton, increasing Z_{eff} .
- Electrons pulled closer, reducing size.

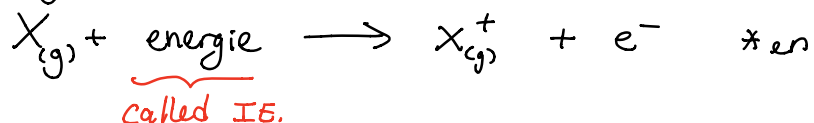
So: Atomic radius decreases as we go from left to right across a period.

② Ionization energy

- Ionization energy (IE): amount of energy needed to REMOVE a valence electron from a ground-state gaseous atom.

LOSS of e^-

- Need energy to overcome attraction between ^{no additional e^- needed to break chem bonds} electrons and protons!



- In the case of multiple valence electrons, there will be multiple IE values (one for each of the valence electrons to remove). More on this after.

- Trend for 1st ionization energy (IE_1)

- Group: \rightarrow Atomic radius \uparrow
 \rightarrow "Easier" (takes less energy) to remove e^- since we are further away from the nucleus

So overall: IE_1 decreases as we go down a group.

- Period: \rightarrow Z_{eff} increases from left to right across period.
 \rightarrow Valence e^- feel the Z_{eff} more strongly, which means we need more energy

So overall: IE_1 increases as we go from left to right across a period.

IMPORTANT: There are exceptions across a period

① From Beryllium to Boron



- The atom is already in a stable state with a full sublevel, requiring more energy to remove one of the 2s electrons.
- Additionally: the 2s electrons feel a stronger attraction to the nucleus (think of Z_{eff}), requiring more energy to overcome the attraction.



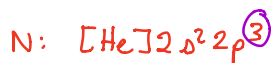
- Removal of the only 2p orbital electron will lead to a more stable state (a full $2s^2$ orbital)

From Magnesium to Aluminium



} similar to Be and B

② From Nitrogen to Oxygen

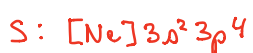
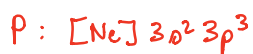


- This electron configuration is already stable since the 2p orbital is half-filled. Removing a 2p electron would "remove" this stability and more energy will be required for the removal.



- Removing a 2p electron will produce a more stable half-filled 2p orbital, which increases the stability of the electron configuration of the atom.
- Additionally: By removing a 2p electron, there will be less repulsion between the 2p electrons.

From Phosphorus to Sulfur



} similar to N and O.

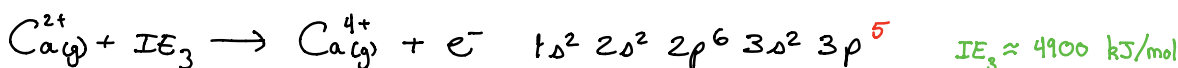
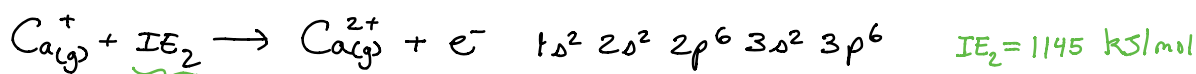
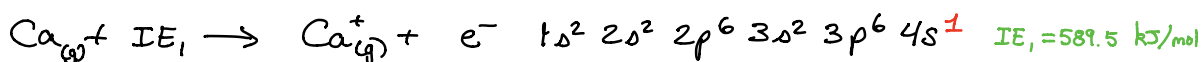
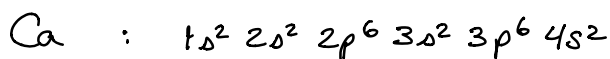
A note on multiple IE values.

- If more than 1 electron, then there are multiple IEs

$$IE_1 < IE_2 < IE_3 < \dots$$

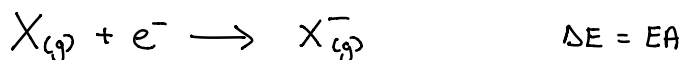
Why does it increase?
- by removing an e^- ,
the other electrons
can feel Z_{eff} more
strongly (get's harder
to remove subsequent e^-).

Look at calcium :



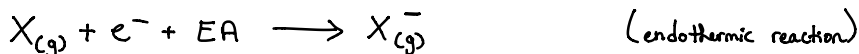
③ Electron Affinity

- Electron affinity (EA): the change in energy produced when an electron is **ADDED** to an atom in the gaseous state.
- Similarly to IE, if multiple electrons are being removed, there are multiple EA values.



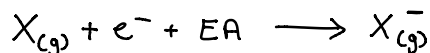
- From table, EA can be either positive or negative.

- IF EA is positive ($EA > 0$):



- We need energy to add the e^- , which means that the atom is less stable (needs energy to keep e^-)

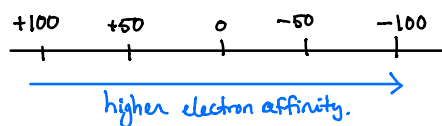
- IF EA is negative ($EA < 0$):



- Energy is released, meaning that adding an electron provides stability.
- In most cases, $EA_1 > 0$ since the e^- is attracted to the nuclear charge.
(first electron affinity)

- EA_2 is always positive. Why? need to absorb energy in order to overcome the electrostatic repulsion experienced when e^- is added.

- An important note:
 - Large negative numbers mean high electron affinity.
 - Small negative numbers and positive numbers mean a low electron affinity.



- Trend for 1st EA (more irregular than IE)

- Group → atomic size increases as we go down a group
 (General trend, it does not always hold) → more e⁻ shielding the valence e⁻.
 → Smaller Z_{eff} , meaning that the electrostatic attraction may not be enough to easily add e⁻ to atom.

So overall: EA₁ decreases as we go down a group

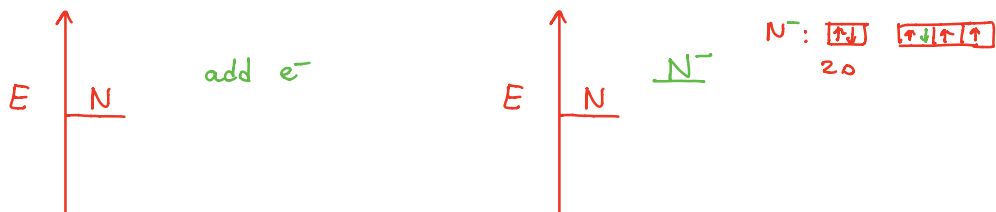
- Period: → increasing Z_{eff} , therefore the electrostatic attraction (between e⁻ and nucleus) enough to facilitate (make easier) the addition of e⁻.
 (General trend, it does not always hold)

• Some exceptions

① Nitrogen. ($EA_1 = +7 \text{ kJ/mol}$)

• Observe the (valence) orbital diagram N: $\boxed{\uparrow\downarrow} \boxed{\uparrow} \boxed{\uparrow} \boxed{\uparrow}$
 2s

- The electron being added must be paired with an unpaired electron in one of the 2p orbitals.
- The sublevel 2p³ is half-full, stabilizing the electron configuration, thus adding an electron to it is not favoured.



② Na ($EA_1 = -52.9 \text{ kJ/mol}$)

• Observe the (valence) orbital diagram Na: $\boxed{\uparrow}$
 3s

- By adding an electron to the 3s orbital, the stability increases since we are filling a full sublevel.
- As such adding an e⁻ is welcomed.

- Recall that when speaking about stability:

- ① Better to have a full level (n)

- ② Better to have a full sublevel (l)

- ③ Better to have a half-full sublevel (l).

2